10.0 Grams Of Argon And 20.0 Grams Of Neon Are Placed In A 1200.0 Ml Container At 25.0 C The Partial

10.0 Grams Of Argon And 20.0 Grams Of Neon Are Placed In A 1200.0 Ml Container At 25.0 C The Partial pressure of each gas and their combined behavior in the container can be analyzed using principles of Dalton's Law of Partial Pressures, ideal gas law, and molar calculations. This article aims to provide a comprehensive understanding of the scenario, including the calculations involved, the concepts of partial pressure, molar quantities, and how these gases interact within a confined space at a given temperature.

Understanding the Scenario

Basic Setup

In the scenario, a container with a volume of 1200.0 milliliters (mL) is filled with two gases: argon (Ar) and neon (Ne). The respective masses are 10.0 grams for argon and 20.0 grams for neon, and the temperature of the system is maintained at 25.0°C.

Relevance of the Data

These data points are crucial for calculating:

- The number of moles of each gas in the container
- The partial pressures of argon and neon
- The total pressure within the container
- The applicability of ideal gas behavior at the given conditions

Key Concepts and Principles

Partial Pressure and Dalton's Law of Partial Pressures

Dalton's Law states that in a mixture of non-reacting gases, each gas exerts a pressure independently as if it occupied the entire volume by itself. The partial pressure of each gas is proportional to its mole fraction in the mixture:

$$
P_{gas} = X_{gas} \times P_{total}
$$

where:

- P_{gas} is the partial pressure of the gas
- X_{gas} is the mole fraction of the gas
- P_{total} is the total pressure exerted by the gas mixture

Ideal Gas Law

The ideal gas law relates pressure, volume, temperature, and amount of gas:

$$
PV = nRT
$$

where:

- P = pressure (atm)
- V = volume (liters)
- n = number of moles
- R = ideal gas constant (0.0821 L·atm/(mol·K))
- T = temperature (Kelvin)

This law helps convert between mass, moles, and pressure.

Calculating Moles of Argon and Neon

Molar Masses of Gases

To convert mass to moles, we need the molar masses:

- Argon (Ar): approximately 39.95 g/mol
- Neon (Ne): approximately 20.18 g/mol

Number of Moles of Argon

$$
n_{Ar} = \frac{\text{mass of Ar}}{\text{molar mass of Ar}} = \frac{10.0\,g}{39.95\,g/mol} \approx 0.250\,mol
$$

Number of Moles of Neon

$$
n_{Ne} = \frac{\text{mass of Ne}}{\text{molar mass of Ne}} = \frac{20.0\,g}{20.18\,g/mol} \approx 0.992\,mol
$$

Converting Volume and Temperature for Gas Calculations

Volume in Liters

Since the volume is given in milliliters, convert it to liters:

$$
V = 1200.0\,mL = 1.200\,L
$$

Temperature in Kelvin

Temperature in Kelvin:

$$
T = 25.0^\circ C + 273.15 = 298.15\,K
$$

Calculating Total Pressure in the Container

Using the ideal gas law for the total amount of gas:

$$
n_{total} = n_{Ar} + n_{Ne} \approx 0.250 + 0.992 = 1.242\,mol
$$

Applying the ideal gas law:

$$
P_{total} = \frac{nRT}{V} = \frac{(1.242\,mol)(0.0821\,L\,atm/(mol\,K))(298.15\,K)}{1.200\,L}
$$

Calculating numerator:

\[

$$
(1.242)(0.0821)(298.15) \approx 30.36
$$

Dividing by volume:

$$
P_{total} \approx \frac{30.36}{1.200} \approx 25.30\,atm
$$

Thus, the total pressure exerted by the gases in the container is approximately 25.30 atm.

Determining Partial Pressures of Argon and Neon

Using Dalton's Law:

$$
X_{gas} = \frac{n_{gas}}{n_{total}}
$$

- Mole fraction of argon:

$$
X_{Ar} = \frac{0.250}{1.242} \approx 0.201
$$

- Mole fraction of neon:

$$
X_{Ne} = \frac{0.992}{1.242} \approx 0.799
$$

Partial pressures:

- Argon:

$$
P_{Ar} = X_{Ar} \times P_{total} \approx 0.201 \times 25.30 \approx 5.09\,atm
$$

- Neon:

$$
P_{Ne} = X_{Ne} \times P_{total} \approx 0.799 \times 25.30 \approx 20.21\,atm
$$

These partial pressures tell us how much each gas contributes to the total pressure within the container.

Implications and Applications

Industrial Uses

Understanding the partial pressures of noble gases like argon and neon is vital in industries such as:

- Lighting (neon lights)
- Welding (argon shielding gases)
- Laser technology (neon lasers)
- Manufacturing processes requiring inert atmospheres

Scientific Research

Accurate calculations of partial pressure are essential for:

- Gas mixture behavior analysis
- Experimental setup calibration
- Safety assessments in handling high-pressure gases

Environmental Considerations

Knowing the behavior of noble gases helps in:

- Atmospheric studies
- Monitoring gas emissions
- Understanding natural gas reservoirs

Limitations and Assumptions

While the calculations above assume ideal behavior, real gases deviate slightly under certain conditions. However, at 25°C and typical laboratory pressures, noble gases like argon and neon behave approximately as ideal gases.

Key assumptions include:

- No chemical reactions occur between gases
- Gases are uniformly mixed
- The container is rigid and maintains constant volume
- Temperature remains constant during the observations

Conclusion

Placing 10.0 grams of argon and 20.0 grams of neon into a 1200.0 mL container at 25.0°C results in a total pressure of approximately 25.30 atm, with partial pressures of about 5.09 atm for argon and 20.21 atm for neon. These calculations demonstrate fundamental principles of gas laws and molar relationships, which are essential in various scientific and industrial applications involving noble gases.

Understanding these concepts enables precise control and utilization of gas mixtures, fostering advancements in technology, research, and environmental management.

Frequently Asked Questions

What is the partial pressure of argon in the container?

Using the ideal gas law, partial pressure of argon can be calculated as P = (nRT)/V. With 10.0 g of argon (molar mass ≈ 39.95 g/mol), R = 0.0821 L·atm/(mol·K), T = 25°C (298 K), and V = 1200 mL (1.2 L), the partial pressure of argon is approximately 8.14 atm.

What is the partial pressure of neon in the container?

Neon's partial pressure is calculated similarly. Moles of neon = 20.0 g / 20.18 g/mol ≈ 0.99 mol. Using P = (nRT)/V, the partial pressure of neon is approximately 16.88 atm.

What is the total pressure exerted in the container?

Total pressure = partial pressure of argon + partial pressure of neon ≈ 8.14 atm + 16.88 atm = 25.02 atm.

How do you determine the mole fraction of argon in the mixture?

Mole fraction of argon (X_Ar) = moles of argon / total moles of gases. Total moles = 10.0 g/39.95 g/mol + 20.0 g/20.18 g/mol ≈ 0.25 mol + 0.99 mol = 1.24 mol. Therefore, X_Ar ≈ 0.25 / 1.24 ≈ 0.20.

What is the mole fraction of neon in the mixture?

Mole fraction of neon (X_Ne) = 0.99 mol / 1.24 mol ≈ 0.80.

How does the partial pressure of each gas change if the temperature increases?

According to the ideal gas law, increasing temperature increases the partial pressure of each gas proportionally, assuming volume and moles remain constant.

What is the significance of partial pressure in this gas mixture?

Partial pressure indicates the contribution of each gas to the total pressure, which is important for understanding gas behavior and calculating individual gas properties within the mixture.

How would the total pressure change if the volume of the container decreased?

Reducing the volume increases the total pressure according to Boyle's Law, since pressure is inversely proportional to volume at constant temperature and moles.

Can Dalton’s Law of Partial Pressures be applied here?

Yes, Dalton’s Law states that the total pressure is the sum of the partial pressures of individual gases, which applies in this scenario.

What assumptions are made when using the ideal gas law for this problem?

Assumptions include that gases behave ideally, interactions between gas particles are negligible, and the gas particles occupy no volume compared to the container volume.

Additional Resources

Understanding the Partial Pressures of Argon and Neon in a Mixture at 25°C

When dealing with gases in a mixture, such as argon (Ar) and neon (Ne), it’s essential to understand how their individual quantities influence the overall behavior of the system—particularly their partial pressures. This detailed analysis explores the scenario where 10.0 grams of argon and 20.0 grams of neon are placed in a 1200.0 mL container at 25.0°C, focusing on the calculations and concepts related to their partial pressures, mole fractions, and the ideal gas law.

Fundamental Concepts Involved

Before diving into calculations, it’s important to review key concepts:

1. Ideal Gas Law

- Expressed as PV = nRT
- Variables:
- P = pressure (atm)
- V = volume (L)
- n = number of moles
- R = universal gas constant (0.082057 L·atm/(mol·K))
- T = temperature (Kelvin)

2. Moles and Molar Mass

- Moles (n) relate mass and molar mass:
- n = mass / molar mass

3. Partial Pressure and Dalton’s Law

- Total pressure (P_{total}) in a mixture is the sum of individual partial pressures:
- $P_{total} = P_{Ar} + P_{Ne}$
- Partial pressure of a component:
- $P_{component} = (n_{component} / n_{total}) P_{total}$
- Mole fraction (X):
- $X_{component} = n_{component} / n_{total}$

Step-by-Step Calculation

Let's analyze the problem methodically:

1. Convert Masses to Moles

Given:
- Mass of Argon = 10.0 g
- Mass of Neon = 20.0 g

Molar masses:
- Argon (Ar) ≈ 39.95 g/mol
- Neon (Ne) ≈ 20.18 g/mol

Calculations:
- n_{Ar} = 10.0 g / 39.95 g/mol ≈ 0.2503 mol
- n_{Ne} = 20.0 g / 20.18 g/mol ≈ 0.9912 mol

2. Determine Total Moles in the Mixture

Total moles:
- $n_{total} = n_{Ar} + n_{Ne} \approx 0.2503 + 0.9912 \approx 1.2415$ mol

3. Convert Container Volume to Liters

Given:
- Volume V = 1200.0 mL

Conversion:
- V = 1200.0 mL × (1 L / 1000 mL) = 1.2 L

4. Convert Temperature to Kelvin

Given:
- T = 25.0°C

Conversion:
- T = 25.0 + 273.15 = 298.15 K

5. Calculate Total Pressure Using Ideal Gas Law

Applying PV = nRT:
- $P_{total} = (n_{total} \times R \times T) / V$

Plugging in values:
- P_{total} = (1.2415 mol × 0.082057 L·atm/(mol·K) × 298.15 K) / 1.2 L

Calculations:
- Numerator: 1.2415 × 0.082057 × 298.15 ≈ 30.33
- Total Pressure: P_{total} ≈ 30.33 / 1.2 ≈ 25.28 atm

Result: The total pressure inside the container is approximately 25.28 atm.

Partial Pressures of Argon and Neon

Now, understanding the individual partial pressures:

1. Mole Fractions

- Mole fraction of Argon:
- $X_{Ar} = n_{Ar} / n_{total} \approx 0.2503 / 1.2415 \approx 0.2017$
- Mole fraction of Neon:
- $X_{Ne} = n_{Ne} / n_{total} \approx 0.9912 / 1.2415 \approx 0.7983$

2. Calculating Partial Pressures

- $P_{Ar} = X_{Ar} \times P_{total} \approx 0.2017 \times 25.28$ atm ≈ 5.10 atm
- $P_{Ne} = X_{Ne} \times P_{total} \approx 0.7983 \times 25.28$ atm ≈ 20.18 atm

Observation: Neon exerts a much higher partial pressure due to its larger molar amount, which aligns with its higher mole fraction.

Implications and Real-World Relevance

Understanding the partial pressures of gases in a mixture plays a critical role in various scientific and industrial processes:

1. Gas Mixture Behavior
- Partial pressures determine diffusion rates, effusion, and mixing behavior.
- Neon, with a higher partial pressure, will dominate the gas phase behavior in terms of partial pressure contribution.

2. Applications in Lighting and Neon Signs
- Neon’s prominence in the mixture influences its luminous properties and color emissions in lighting applications.

3. Industrial Gas Storage and Transportation

- Accurate partial pressure calculations ensure safety and efficiency when handling compressed gases.

4. Scientific Experiments and Calibration

- Precise control over partial pressures is crucial for experiments involving gas reactions, spectrometry, and other analytical techniques.

Additional Considerations and Assumptions

While the calculations above assume ideal behavior, real gases can deviate slightly:

1. Deviations from Ideal Gas Law

- At high pressures or low temperatures, gases may deviate from ideality.
- Intermolecular forces and finite molecular sizes can influence pressure.

2. Gas Purity

- Assumes gases are pure; impurities can alter the partial pressures.

3. Container Conditions

- No leaks, uniform temperature, and isotropic conditions are presumed.

4. Measurement Accuracy

- Precise measurements of mass, volume, and temperature are essential for accurate calculations.

Extended Analysis: Impact of Changing Variables

Understanding how variations in parameters influence the system:

1. Effect of Temperature Increase

- Raising temperature increases total pressure proportionally.
- Partial pressures of each gas would also rise, assuming constant moles and volume.

2. Effect of Volume Reduction

- Decreasing volume raises total pressure.
- Partial pressures increase proportionally, affecting gas behavior and potential reactions.

3. Changes in Gas Amounts

- Adding more of either gas shifts mole fractions and partial pressures.
- For example, increasing neon amount would raise its partial pressure more significantly.

Conclusion

This comprehensive analysis underscores the importance of fundamental principles like the ideal gas law, mole fraction calculations, and Dalton's law in understanding gas mixtures. In the scenario of 10.0 grams of argon and 20.0 grams of neon in a 1200.0 mL container at 25.0°C, the total pressure is approximately 25.28 atm, with argon contributing roughly 5.10 atm and neon about 20.18 atm as partial pressures.

Such calculations are vital in scientific research, industrial processes, and practical applications where precise control and understanding of gas behaviors are essential. They illustrate how even simple mass and volume measurements can lead to deep insights into the microscopic and macroscopic properties of gas mixtures, bridging the gap between theoretical principles and real-world applications.

[10 0 Grams Of Argon And 20 0 Grams Of Neon Are Placed In A 1200 0 Ml Container At 25 0 C The Partial]

10 0 Grams Of Argon And 20 0 Grams Of Neon Are Placed In A 1200 0 Ml Container At 25 0 C The Partial

Related Articles

- 1. Which Pair Of Atoms Will Share Electrons When A Bond Is Formed Between Them? Explain

Why.A. K And

- A Partnership Began Its First Year Of Operations With The Following Capital Balances: Young, Capital
- REASONING Do You Always Need To Regroup Whenmultiplying 3-digit Or 4-digit Numbers? Why Or Why Not?

Back to Home